Chemistry EOY Notes

Acids and Bases, Redox Reactions, Chemical Energetics, Reaction Kinetics

Ryan Sam (Class 3P)

Content Page:

<u>Acids and Bases</u>:

- Definition of Acid: an acid is a substance that ionizes in water to produce hydrogen ions
- Definition of Alkali: an alkali is a soluble base that ionizes in water to produce hydroxide ions

Common Acids:

Common Name	Chemical Name	Formula	Description
Hydrochloric Acid	Aqueous Hydrogen Chloride	HCl	Strong acid, monobasic
Nitric Acid	Aqueous Hydrogen Nitrate	$HNO3$	Strong acid, monobasic
Sulfuric Acid	Aqueous Hydrogen Sulfate	$H2SO4$	Strong acid, dibasic
Carbonic Acid	Aqueous Hydrogen Carbonate	$H2CO3$	Weak acid, dibasic
Acetic acid (Vinegar)	Aqueous Ethanoic Acid	$CH3COOH$	Weak acid, monobasic

Common Alkalis:

Common Name	Chemical Name	Formula	Description
Caustic Soda	Sodium hydroxide	$NaOH$	Strong alkali
Caustic Potash	Potassium hydroxide	KOH	Strong alkali
Slaked lime	Calcium Hydroxide	$Ca(OH)2$	Strong alkali (but only slightly soluble in water)
Ammonia solution	Aqueous Ammonia	$NH3$	Weak alkali

- Aqueous ammonia is formed by dissolving $NH3$ in water. It is to be written as $NH3$ (aq). It is alkaline because some ammonia molecules accept protons from water to produce $NH4+$ and $OH-$ ions

<u>Chemistry of Acids:</u>

1. Acid reacts with Base to form Salt and Water (Neutralization)
2. Acid reacts with Carbonate to form Salt, Water and Carbon Dioxide
3. Acid reacts with (some) Metal to form Salt and Hydrogen
4. Acid reacts with Sulfite to form Salt, Water and Sulfur Dioxide

<u>Chemistry of Bases:</u>

1. Base reacts with Acid to form Salt and Water (Neutralization)
2. Base reacts with Ammonium salt to form Salt, Water and Ammonia
3. Alkali reacts with (some) Salt solutions to form Salt and insoluble Hydroxides (precipitation reactions)

<u>Strength and Basicity of Acids:</u>

Strength:

- Extent of dissociation/ionization
- A strong acid has acid molecules that are easily (or completely) ionized in an aqueous solution
- A weak acid has molecules that are partially ionized in aqueous solution, with a high proportion of the acid remaining in the undissociated form
- To represent a weak acid's reaction with water, a reversible arrow is used

Concentration:

- Measurement of the amount of solute in a solvent
- Represented by gdm-3 or moldm-3 (Molarity)
- The industrial benchmark for a concentrated acid is 2 moldm-3

Basicity:

- Dependent on the number of H atoms in a molecule that are able to ionize to form $H+$ ions
- Hydrochloric acid is able to provide 1 hydrogen ion per molecule (monobasic)
- Sulfuric acid is dibasic while phosphoric acid is tribasic
- Not all the hydrogen atoms in a molecule ionizes to form $H+$ ions (in ethanoic acid, $CH3COOH$, only the H in the carboxyl functional group is dissociated)

- The basicity of an acid does not represent the strength of an acid. The number of hydrogen ions liberated per molecule of acid does not determine its strength. A dibasic or tribasic acid is not necessarily stronger than a monobasic acid. (Nitric acid and hydrochloric acid are strong monobasic acids while carbonic acid is a weak dibasic acid)

Conditions for dissolving an acid:

- Requires water (or a polar solvent capable of dissociating the acid)
- The acid must dissolve in the water and dissociation must occur to form H+ ions (before it displays its acidic properties)
- HA + H2O → H3O+ + A- (where the acid molecule has been ionized, with A representing a molecule bonded to the hydrogen ion)
- In an aqueous state, the ions are surrounded by water molecules (and the hydrogen ions cause a solution to be acidic and exhibit acidic properties)
- Important Experiment: when hydrogen chloride gas or citric acid solid is dissolve in nonpolar solvents such as toluene they do not exhibit acidic properties as their molecules do not ionize (dissociate). Thus it is the presence of the H+ ions in aqueous solutions that give these acids the acidic properties

Writing Formulae of Salts (Neutralization):

1. When an acid reacts with a base, the formula of the salt formed is derived by replacing the H+ of the acid and the metal ion of the base
2. This replacement of the H+ must obey the rule of balancing charges (when HCl reacts with MgO, only MgCl2 can be formed, since one Mg2+ replaces 2 H+)
3. For dibasic and tribasic acids, since there are 2H+ that can be replaced, 2 possible salts can form
    - When H2SO4 reacts with NaOH, if one H+ is replaced, the salt formed is NaHSO4. If two H+ are replaced, then Na2SO4 is formed.

Reactions of Neutralization:

- 4 main types of observation:
    - Production of a gas
    - Production of a solid (precipitate)
    - Dissolving of a solid to form a solution
    - Colour change of pH indicator

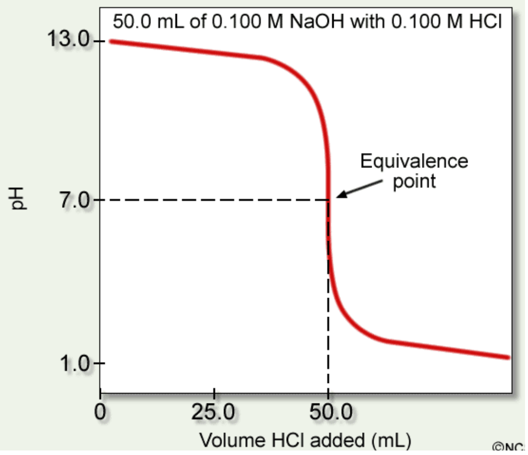
pH of a Solution:

- pH=power of hydrogen (logarithmic scale)
- pH of a solution measures the extent of alkalinity or acidity of a solution
- pH of a solution is given as the negative logarithm to base ten of the molar hydrogen ion concentration (pH=-lg(H+))
- A change in one pH unit changes the hydrogen ion concentration by a factor of ten
- pH values can be above 14 and lower than 0: only in cases of high concentration of acids/bases
- Important: an aqueous solution always contains H+ and OH- ions, due to the small extent of partial dissociation of water molecules (H2O → H+ and OH-)
    - If a solution is acidic, [H+] > [OH-] and the pH < 7
    - If a solution is neutral, [H+] = [OH-] and the pH=7
    - If a solution is alkali, [H+] < [OH-] and the pH > 7
- pH of a solution can be measured by:
    - Chemical indicators which show different colours at different pH values (Universal Indicator and litmus paper)
    - pH meter/sensor: an electrical device more accurate than indicators and must be dipped into the solution

Acid-Base Titration Curve:

- Acid-Base Titration: Used to determine the concentration of an acid or base. This is done by exactly neutralizing the acid/base with a base or an acid with known concentration
- When acid and base are mixed in exactly the right proportions for neutralization, the equivalence point is reached.
- When the indicator changes colour, this is the “end point” of a titration
- NOTE: whether the acid or the base was originally in the beaker (affects the shape of the curve)

Strong Acid-Strong Base Titration:

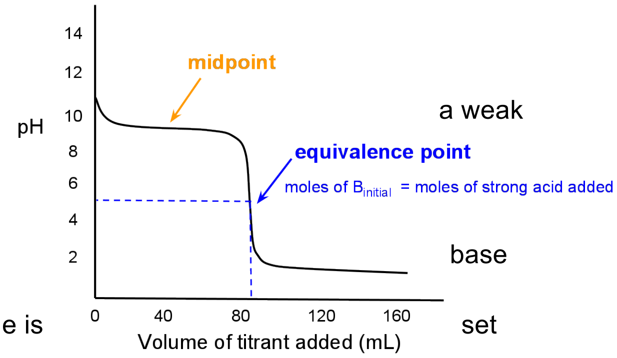


- The pH decreases slowly at first (pH only falls a small amount until it is close to the equivalence point) because of high concentration of OH- ions ([OH-] > [H+])
- Gives rise to a neutral salt (since the equivalence point is at pH7)
- Each drop of HCl adds in a high concentration of H+ ions

Explanation:

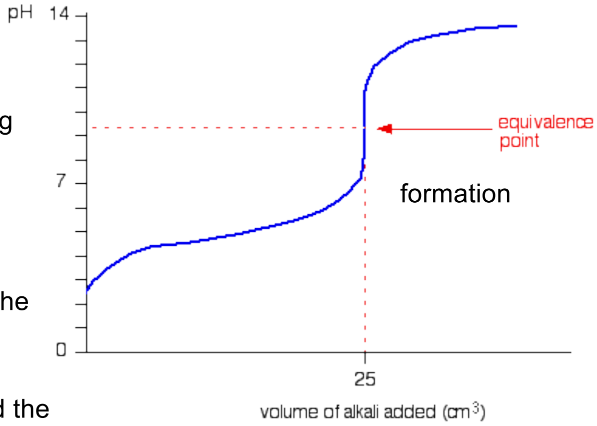
- Since the pH scale is on a logarithm, increasing the pH from 4 to 10 is easier than increasing the pH from 1 to 2. At pH 1, [H+] is 0.1 moldm-3, at pH 2 [H+] is 0.01, while at 3 it is 0.001
- This makes it comparatively easier to increase the pH after a certain point, and it always cuts through the equivalence point

Strong Acid-Weak Base: (probable EOY question)



- Example: aqueous ammonia and hydrochloric acid
- 2 sharp drops in pH
- Start offs with a pH lower than 13/14 since it is only a weak base (partial dissociation)
- Ends off with a pH of 1 since a strong acid is added (complete H+ dissociation)
- When the strong acid is initially added to the weak base the fall in pH is drastic. But very quickly a buffer solution of aqueous ammonia and ammonium chloride is set up, so the pH falls slowly
- The equivalence point is at a pH less than 7,because the salt formed (ammonium chloride) is acidic (strong acid-weak base titration usually have an equivalence point below 7)
- Smaller gradient of decrease compared to strong acid-strong base titration

Weak Acid-Strong Base:



- The starting pH should be greater than 1-2, since a weak acid is used (partial H+ dissociation)
- The ending pH should be close to 13-14, since a strong base is used (where the [OH-] > [H+])
- The equivalence point is greater than 7 (around 9-10): formation of a basic salt/solution

Weak Acid-Weak Base:

- Not often titrated since the colour change shown with the indicator is often quick, and therefore difficult for the observer to see the change of colour
- The graph does not have a sharp plunge in the pH and the equivalence point is difficult to determine

Types of Oxides:

- If there is a reaction, a salt must be formed

- Hydroxides behave the same way as metal oxides (sodium hydroxide, calcium hydroxide and copper (II) hydroxide are basic while aluminium hydroxide, lead (II) hydroxide and zinc hydroxide are amphoteric)

Metal Oxides:

- Basic Oxides: react with acids to form salts
    - Including sodium oxide (Na_2O), calcium oxide (CaO) or copper (II) oxide (CuO)
    - Mainly Group I and II oxides (with transition metals)
    - (i) $CaO + 2HCl \rightarrow CaCl_2 + H_2O$ (where the Ca^{2+} ions replace the H^+ ions)
- Amphoteric Oxides: react with both acids and bases to form salts
    - Only the "ZAP" (Zinc, Aluminium, Lead) oxides are amphoteric
    - Including zinc oxide (ZnO), Aluminium oxide (Al_2O_3), Lead (II) oxide (PbO)
    - (i) $Al_2O_3+6HCl \rightarrow 2AlCl_3+3H_2O$
    - (ii) $Al_2O_3+ 2NaOH \rightarrow 2NaAlO_2$ (sodium aluminate, a soluble complex salt) $+H_2O$

Non-Metal Oxides:

- Acidic Oxides: react with bases to form salts
    - Including carbon dioxide (CO_2), sulfur dioxide (SO_2) and nitrogen dioxide (NO_2)
    - Many of them are major air pollutants
    - Usually non-metals (that react with bases)
    - (i) $CO_2+Ca(OH)_2 \rightarrow CaCO_3+H_2O$
- Neutral Oxides: do not react with acids or bases

<u>Solubility Table:</u>

- Rule 1: Most chlorides, bromides and iodides are soluble in water. The two common exceptions are silver and lead (II) chloride/bromide/iodide
- Rule 2: Most sulfates are soluble. The three common exceptions are barium sulfate, calcium sulfate and lead (II) sulfate
- Rule 3: All carbonates are insoluble in water except ammonium carbonate and group I carbonates
- Rule 4: All oxides and hydroxides are insoluble in water except Group (I) oxides/hydroxides. (Calcium oxide and hydroxide are sparingly soluble)
- Rule 5: All group I and ammonium compounds are soluble in water
- Rule 6: All nitrates are soluble in water

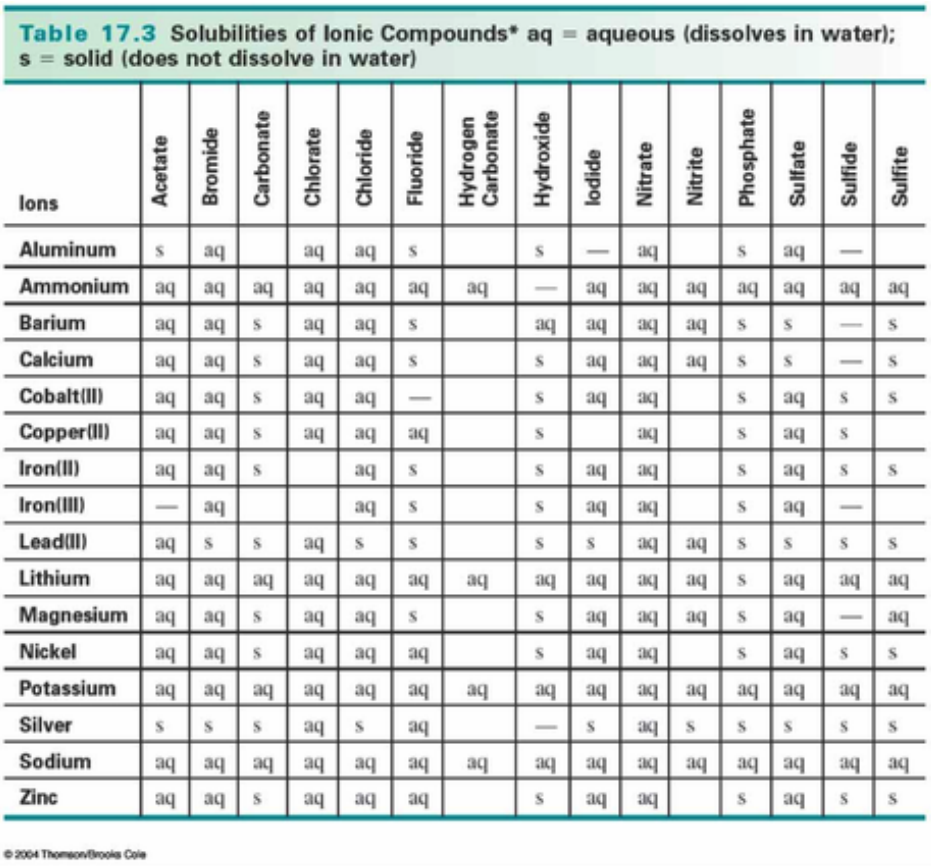


Table 17.3 Solubilities of Ionic Compounds* aq = aqueous (dissolves in water); s = solid (does not dissolve in water)

Ions	Acetate	Bromide	Carbonate	Chlorate	Chloride	Fluoride	Hydrogen Carbonate	Hydroxide	Iodide	Nitrate	Nitrite	Phosphate	Sulfate	Sulfide	Sulfite
Aluminum	s	aq		aq	aq	s		s	—	aq		s	aq	—	
Ammonium	aq	aq	aq	aq	aq	aq	aq	—	aq	aq	aq	aq	aq	aq	aq
Barium	aq	aq	s	aq	aq	s		aq	aq	aq	aq	s	s	—	s
Calcium	aq	aq	s	aq	aq	s		s	aq	aq	aq	s	s	—	s
Cobalt(II)	aq	aq	s	aq	aq	—		s	aq	aq		s	aq	s	s
Copper(II)	aq	aq	s	aq	aq	aq		s		aq		s	aq	s	
Iron(II)	aq	aq	s		aq	s		s	aq	aq		s	aq	s	s
Iron(III)	—	aq			aq	s		s	aq	aq		s	aq	—	
Lead(II)	aq	s	s	aq	s	s		s	s	aq	aq	s	s	s	s
Lithium	aq	aq	aq	aq	aq	aq	aq	aq	aq	aq	aq	s	aq	aq	aq
Magnesium	aq	aq	s	aq	aq	s		s	aq	aq	aq	s	aq	—	aq
Nickel	aq	aq	s	aq	aq	aq		s	aq	aq		s	aq	s	s
Potassium	aq	aq	aq	aq	aq	aq	aq	aq	aq	aq	aq	aq	aq	aq	aq
Silver	s	s	s	aq	s	aq		—	s	aq	s	s	s	s	s
Sodium	aq	aq	aq	aq	aq	aq	aq	aq	aq	aq	aq	aq	aq	aq	aq
Zinc	aq	aq	s	aq	aq	aq		s	aq	aq		s	aq	s	s

Precipitation Reactions:

- Precipitation is the formation of a solid when 2 aqueous solutions are mixed. An insoluble substance would appear as a precipitate when it is formed.
- To form the precipitate, one of the reactants used must contain the cation (positive ion) of the ppt and the other reaction must contain the anion (negative ion) of the ppt. Both ions must be in aqueous solutions in the reactants
- A precipitate must be formed by mixing 2 aqueous solutions. It should not be formed by mixing a solid with a solution.
  - (i) PbO (s) + H_2SO_4 (aq) → $PbSO_4$ (s) + H_2O (l)
  - This does not work since the solid $PbSO_4$ would form as a layer on the solid PbO and cover it, thus stopping the reaction prematurely as the acid could no longer react with the remaining PbO that is covered by the $PbSO_4$.
- Collection of precipitate:
  - Step 1: Filter the mixture
  - Step 2: Wash the residue on the filter paper with distilled water
  - Step 3: Dry the solid by pressing it between sheets of filter paper or in a low-heat oven
- Note: Aqueous lead (II) nitrate and dilute hydrochloric acid
  - The insoluble salt formed will be lead (II) chloride
  - *In this case, hydrochloric acid does not act as an acid, but as a salt
  - There is no acid-salt reaction, so the salt reacts with another salt to form an insoluble precipitate and nitric acid
- Formula of insoluble precipitate PbI_2
  - Reactant 1 in aqueous form: $Pb(NO_3)_2$ (try to use nitrates since they are soluble)
  - Reactant 2 in aqueous form: KI (try to use a Group I metal since they are soluble)

Ionic Reactions:

- When substances reach, not all the particles will take part in the reaction
- An ionic reaction specifically shows which atoms/ions/molecules took part in the reaction, through eliminating the spectator ions
- Spectator ions are free ions present at the start and at the end of a reaction (they remain unchanged by the reaction)
- The basic steps:
  1) Write the balanced equation (Pb(NO3)2 + 2KI → PbI2 + 2KNO3)
  2) Put in the state symbols (Pb(NO3)2 (aq) +2KI (aq) → PbI2 (s) + 2KNO3 (aq))
     - Water-soluble substances are in the aqueous state
  3) Write out the free ions in the aqueous solution
     Pb2+ (aq) + 2NO3- (aq) + 2K+ (aq) +2I- (aq) → PbI2 (s) + 2K+ (aq) +2NO3- (aq)
     - Substances in solid, liquid or gaseous state do not form free ions
  4) Cancel away the corresponding spectator ions on the left hand side and right hand side of the equation (2K+ , 2NO3-)
  5) Write down the uncancelled formula gives the ionic equation
     Pb2+ (aq) + 2I- (aq) → PbI2 (s)
- State symbols must be included in the final ionic solution (Step 5)
- The number of each type of atom on both sides of the equation are cancelled, the net charge on both sides are also balanced
- The ionic reaction shows that the formation of lead (II) iodide requires aqueous lead (II) and aqueous iodide ions
  - Pb (NO3)2 + 2NaI → PbI2 + 2NaNO3
  - Pb (CH3COO)2 + 2KI → PbI2 + 2KCH3COO
  - Pb (CH3COO)2+ 2NaI → PbI2 + 2NaCH3COO
  - All of them will have the same ionic reaction
- All neutralization reactions have the same ionic reactions (the soluble salt formed in the reaction remains as ions in the solution, while water becomes a liquid)
  H+ (aq) + OH- (aq) → H2O (l)
- The ionic equation of a precipitation involves only those ions of the precipitate
- Acid metal reaction: the metal usually displaces the H+ ion of the acid
  Zn (s) + H2SO4 (aq) → ZnSO4 (aq) +H2 (g)
  Zn (s) + 2H+ (aq) → Zn2+ (aq) + H2 (g)
- Acid and a soluble carbonate (the hydrogen ion and the carbonate form water and carbon dioxide)
  2HNO3 (aq) +Na2CO3 (aq) → 2Na2NO3 (aq) + H2O (l) + CO2 (g)
  2H+ (aq) + CO3(2-) (aq) → H2O (l) + CO2 (g)
- Displacement reaction (the more reactive metal gains electrons from the less reactive metal)
  Zn (s) + CuSO4 (aq) → ZnSO4 (aq) + Cu (s)
  Zn (s) + Cu2+ (aq) → Zn 2+ (aq) + Cu (s)

Additional Material:

Colour of Precipitates:

- AgCl; BaSO4; PbSO4 (White)
- PbI2; PbCrO4 (Yellow)
- CuO (black)
- Ag2CrO4 (Brown/Red) where Cr has an oxidation state of +6
- Al2CrO4 (Yellow/Brown) where Cr has an oxidation state of +2

Acid Rain:

- Involves methods that acids in the atmosphere fall onto the Earth
- Wet deposition: acidic rain, fog and snow

  - Acidic water flows over and through the ground, it affects a variety of plants and animals
  - Strength of its effects depend on the acidity of the water, the chemistry and buffering capacity of the soils, and the types of living organism
  - “Normal” rain is weakly acidic due to the natural presence of small amounts of acidic gases in the atmosphere, but acid rain has a pH of less than 4
- Dry deposition: acidic gases and particles
  - The wind blows these acidic particles and gases onto buildings, cars, homes and trees.
  - Dry deposited gases and particles can be washed from trees and other surfaces by rainstorms: run off water adds to the acids in the acid rain
- Sources of Pollutants
  - Sulfur Dioxide; produced by combustion of sulfur containing fossil fuels in power stations and motor vehicles
  - Oxides of nitrogen are produced in car engine: combined under high temperature to form nitric oxide which is converted to NO2 by reacting with oxygen
- Effect of Pollutants
  - SO2 and oxides of nitrogen damage the lungs, causing bronchitis
  - In the atmosphere they combine with water and oxygen to from nitric acid and sulfuric acid (acid rain)
  - Acid rain damages metal bridges and stone buildings (corrosive) and lowers the pH of water bodies, killing living organisms
- Minimization of Threats
  - Remove sulfur from fossil fuels before they are burnt
  - Remove So2 from waste gases after fossil fuel combustion by flue gas desulfurization (add calcium carbonate: $CaCO3+SO2 \rightarrow CaSO3+CO2$
  - Minimize NO2 pollution by fitting cars with catalytic converters: $2NO + 2CO \rightarrow N2 + 2CO2$ (but carbon monoxide is toxic and lethal as well)

Applications of Neutralization:

1) Controlling the pH of soil
   - Most plants will not grow well if the soil is too acidic
   - Reduce acidity of soil: add calcium oxide or calcium hydroxide (insoluble in water): cannot be washed away by rainwater
2) Treatment of Indigestion
   - Overeating causes the stomach to produce too much hydrochloric acid (relieve pain by consuming mild alkali or a carbonate)
3) Treatment of Insect Stings
   - Pain from a bee sting is due to an acid injected into the skin (neutralized by applying baking soda)
   - Wasp sting (alkaline) is treated with vinegar or lemon juice
4) Treatment of Industrial Wastewater or Waste gases
   - Wastewater can be highly acidic or alkaline (treat wastewater to between pH 6-9 before discharge into the sewage system)
   - Waste gas must undergo flue gas desulfurization
5) Egg dyeing
   - Egg shell is 95% calcium carbonate crystals (covered with cuticle rich in proteins)
   - Dye molecules attach to egg shell: acidic groups (carboxylate or sulfonate SO3H) must from bonds with the protein molecules of the cuticle

Redox Reactions:

- Oxidation and reaction occurs simultaneously in a complementary manner
- One substance is oxidized and another is reduced
- Involves the transfer of electrons between molecules/atoms/substances
- The number of electrons lost by one species always equals to the number of electrons gained by another (it is not lost from the system)

Definitions and terms:

- The substance that oxidizes other atoms/molecules/ions is the oxidizing agent (but is itself reduced)
- The acceptor of the electrons is the oxidizing agent
- The substance that reduces other atoms/molecules/ions is the reducing agent (but is itself oxidized)
- The donor of the electrons is the reducing agent
- A substance is oxidized when a reactant loses electrons during the reaction (OIL: oxidation=loss)
- A substance is reduced when a reactant gains electrons during the reaction (RIG: reduction=gain)
- An element is oxidized when its oxidation state increases
- An element is reduced when its oxidation state decreases
- Disproportionation reaction: the same reactant undergoes both oxidation and reduction at the same time
  - $Cu2O + H2SO4 \rightarrow CuSO4 + Cu (s) + H2O$
  - The Cu in Cu2 has an oxidation state of +1, while Cu in CuSO4 has an oxidation state of +2, and Cu in Cu has an oxidation state of 0

Half Equations:

- Reactions involving a transfer of electrons can be viewed as two halves
- For NaCl (where $2Na+Cl2 \rightarrow 2NaCl$)
  - $2Na \rightarrow 2Na+ + 2e-$
  - $Cl2 + 2e- \rightarrow 2Cl-$

1) Check that the number of electron lost and gain are the same (if not, multiply by a suitable number)
2) Put the left hand side formulae together with the right hand side formulae and cancel away the electrons on both sides to combine the ions

Oxidation States and Oxidation Numbers:

- A number given to an element preceded by + or – sign
- The number is the charge of the atom of the element would have if it exists as an ion in a compound (even if the compound is covalently bonded)

Assigning Oxidation Numbers:

1) Elements that are uncombined with other elements is assigned an oxidation number of 0
2) In simple monoatomic ions, the oxidation number is simply the charge on the ion
3) In polyatomic ions, the sum of the oxidation numbers equals to the charge on the ion
4) The oxidation number of hydrogen in all its compounds is +1, except in metal hydrides where its oxidation number is -1
5) The oxidation number of oxygen in all its compounds is -2, except in peroxides where its oxidation state is -1
6) The oxidation number of Group I elements in their compounds is +1, for Group II elements in their compounds is +2, and for aluminium in its compounds is +3
7) There are many oxidation numbers for Group VII elements in their compounds but the usual one is -1
   - In Cl2O, the oxidation state of Cl is +1

    - For two halogens bonded together, the more electronegative molecule will gain the electron and be -1, the other will be +1 (Cl-Br where Cl is -1 and Br is +1)

8) The sum of the oxidation numbers of all the atoms in a molecule or a compound is 0

Reducing and Oxidizing Agents:

- Reactants that undergo reduction or oxidation with observable colour changes
- Oxidizing agents are used to test for reducing agents, and reducing agents are used to test for oxidizing agents

Potassium Manganate (oxidizing agent):

	MnO_4^- manganate (VII) ion	Mn^{2+} manganese ion
Oxidation Sate of Mn	+7	+2
Colour	Purple	Colourless

- Used to test for reducing agent (for an unknown substance)
- The potassium manganate must be acidified in order to speed up the reaction
- Strongly oxidizing

Potassium Dichromate (oxidizing agent)

	$Cr_2O_7^{2-}$ dichromate (VI) ion	Cr^{3+} chromium (III) ion
Oxidation state of Cr	+6	+3
Colour	Orange	Green

- Used to test for reducing agent (for an unknown substance)
- The potassium dichromate is acidified to speed up the reaction
- Strongly oxidizing

Potassium iodide (reducing agent)

	I^- iodide ion	I_2 iodine
Oxidation State of I	-1	0
Colour	Colourless	Purple/Brown

- Used to test for oxidizing agent (for an unknown substance)

List of Common Oxidizing Agents	List of Common Reducing Agents
▪ O_2 ▪ Cl_2, F_2 and Group 7 elements ▪ $KMnO_4$ ▪ $K_2Cr_2O_7$ ▪ H_2O_2 (when oxygen gas is released) ▪ Fe^{3+}, Pb^{4+} or transition metal ions with higher oxidation states compared to ions of the same element (as Fe^{2+} and Pb^{2+} have comparatively lower oxidation states)	▪ H_2O_2 (will not release oxygen gas) ▪ I^- or all iodides ▪ Cl^- or all chlorides (and all halides) ▪ Most metals (in their neutral state) ▪ CO ▪ SO_2 ▪ Sulphites, eg. Na_2SO_3 (to become SO_4) ▪ Thiosulphates, eg. $Na_2S_2O_3$ (to become SO_4)

Other Useful Colour Changes

Redox Reagents	Redox Property	Colours
I_2, iodine in solution	Oxidizing	Brown; red-brown solution
I_2, iodine crystals	Oxidizing	Violet solids
I_2, iodine vapour	Oxidizing	Violet gas
I_2, iodine precipitate	Oxidizing	Black precipitate
I^-, iodide solution	Reducing	Colourless solution except in Yellow solid in AgI, PbI_2

Fe2+ (aq), Iron (II) solution	Reducing	Light green solution
Fe3+ (aq), Iron (III) solution	Oxidizing	Yellow or Brown Solution
Fe2+ (s) ppt, Iron (II) ppt	Reducing	Dirty Green ppt
Fe3+ (s) ppt, Iron (III) ppt	Oxidizing	Red-brown ppt

<u>How to answer Redox Reactions questions:</u>

- To show that a Redox Reaction has taken place:
    1) X is reduced and the oxidation state of X has decreased from __ in Y to __ in Z
    2) A is oxidized and the oxidation state of A has increased from __ in B to __ in C

Fun Facts:

- Anti-oxidants slow down oxidation (for many berries): they react with oxygen so that the oxygen does not react with the flesh of the fruits
- Less fun fact: oxygen can actually kill humans since it oxidizes the internal body cells

Chemical Energetics:

Types of Energy Conversions:

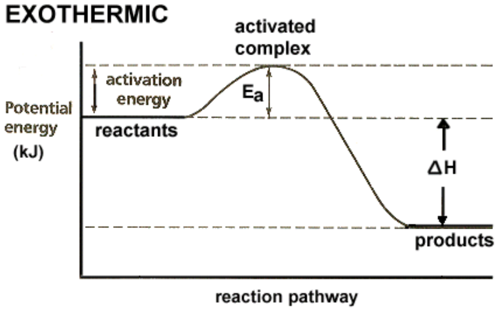
1) Chemical reactions that involve a heat or light change (exothermic and endothermic reactions)
2) Chemical reactions that involve an electrical change (in cells and batteries, in electrolysis)
3) Nuclear reactions (nuclear fission)

Exothermic Reaction and Endothermic Reaction:

- Measures the energy in the particles, lost and gained to the surrounding
- Spontaneous action of bond-formation and bond-breaking

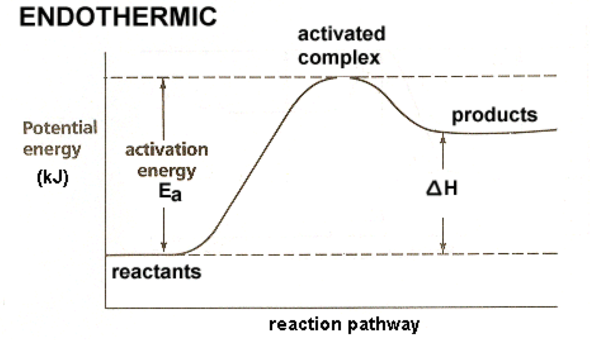
Exothermic reaction:

- Energy (heat) is given out
- Causing a temperature rise in the surroundings
- The energy content of the products is lower than the energy content of the reactants thus the ΔH (enthalpy change or heat of reaction) is a negative value
- The total energy absorbed by the bond-breaking is less than that given out by bond-making
- Includes: combustion, respiration, neutralization, freezing, condensation, and diluting concentrated sulfuric acid
- When a substance burns heat is given out. When a small amount of the substance produces a lot of heat (which is stored in the bonds), then that substance is useful as fuel. Fuels include fossil fuels, wood and hydrogen (fossil fuels store C-C bonds in the hydrocarbon chains)



Endothermic reaction:

- Energy (heat) is taken in
- Temperature decrease in the surroundings
- Products have a higher energy content than the reactants do, thus the ΔH value is positive
- The total energy absorbed by the bond-breaking is more than that given out by bond-making
- Includes: thermal decomposition, photosynthesis, melting and boiling (Photosynthesis is an endothermic reaction that takes in light energy instead of heat)
- Thermal decomposition: heating of limestone
  $CaCO_3 \rightarrow CaO + CO_2 \quad \Delta H = +178kJ$
- Photosynthesis:
  $6CO_2 + 6H_2O \rightarrow C_6H_{12}O_6 + O_2 \quad \Delta H = +2816kJ$
- NOTE: the positive sign in front of ΔH value must be written



Bond Formation and Bond Making:

- During a chemical reaction, bonds are broken and formed
  H_2 (g) +Cl_2 (g) → 2HCl (g)
- The H-H bond and the Cl-Cl bond must first be broken (the bond breaking process is endothermic and requires heat)
- Each H atom forms a bond with each Cl atom (this bond making process is exothermic and gives out heat)

Energy level diagram:

- Activation Energy: energy required to start the chemical reaction
  - Mixing iron and sulfur powder together has a high activation energy. To start the reaction, heat is required (from an external energy source)
  - Magnesium ribbon and acid react readily because there is a low activation energy to start the reaction
- This is usually the strength of the bonds (and the amount of energy required to break them)
- Also known as: energy barrier
- If the reaction requires more heat for activation energy, a catalyst is required to add energy (so that the activation energy is reached for the reaction to occur)
- Bond energy is the energy absorbed in breaking a covalent bond
- ΔH = total energy of bond breaking – total energy of bond making

Other Notes:

- When two elements combine to form a compound and heat is given off, such a compound is an exothermic compound. It has a low energy content and is energetically stable with respect to its elements, thus the bonds between its atoms are strong
- An exothermic compound requires a lot of heat before it will decompose

From $2H_2$ (g) + O_2(g) → $2H_2O$ (l) $\Delta H=-575kJ$

1) Water is an exothermic compound, it is energetically more stable than hydrogen gas
2) The H-O bond in water is stronger than the H-H bond in hydrogen
3) It requires 287.5kJ of energy to decompose 1 mole of water molecules (into its constituent components of hydrogen and oxygen)

Reaction Kinetics: Rate of Reaction

- Involves the rate of reaction (how fast it occurs)
- Energetics refers to how much energy is produced or required

Explanation:

- For a chemical reaction to occur, there must be collisions between the reactant particles, and only a small fraction of the collisions result in a reaction, these are effective collisions
- The rate of reaction depends on the frequency of the effective collisions between reactant particles

Activation Energy:

- All colliding particles require a minimum amount of energy before a reaction can occur (Activation Energy): below this level, a collision cannot occur to produce a reaction
- When the activation energy is high, the reaction is slow, since only a few of the collisions give sufficient energy to produce a reaction (needs catalyst)
    - Formation of water (will only take place if energy is supplied in the form of heat, light or electrical energy
- When the activation energy is low, the reaction is fast, since many colliding particles have energies equal to or greater than the activation energy
    - $2NO + O2 = 2NO2$ (formation of nitrous oxide): can take place spontaneously at room temperature (lightning produces heat and can help in speeding up such reactions)
    - Catalysts make reactions more energetically efficient, and lower the activation energy
    - Catalysts provide a new pathway for reactions (with a lower activation energy) but the original pathway still remains and molecules can use either pathway (unchanged ΔH)
    - Activation Energy: requires the breaking of bonds so new ones are formed (a minimum amount of energy is required at the start of the reaction)

Factors of Rate of Reaction:

I. Increase the temperature
- At a higher temperature, the particles have more kinetic energy and they move faster, colliding with each other harder, thus increasing the frequency of effective collisions
- An increase in the number of reactant particles having the activation energy (derived from heat)
- A rise in 10 degrees Celsius will help in doubling the rate of reaction (for most)

II. Increase the concentration of reactants (only for liquid/aqueous reactants)
- At a higher concentration, there are more particles per unit volume, thus the particles are closer to each other and the frequency of effective collision increases, increasing the reaction rate
- Increasing the concentration of a reactant would also result in the increase in the amount of products (which does not occur for other factors)

III. Increase the pressure (only for gaseous reactants)
- At a higher pressure for gaseous reactants, the particles are in closer proximity and collide more frequently, increasing the reaction rate
- Pressure is used to increase the speed of cooking in a pressure cooker (which increases the number of steam molecules in contact with the food, compared to when steam and water are used at normal atmospheric pressure)

IV. Increase the surface area of reactants (only for solid reactants)
- An increased surface area means that the area of contact between the reactants increases thus the number of effective collisions per unit time also increases
- Small pieces of solid have a larger total surface area than large pieces of the same mass (surface area to volume ratio): stirring will cause larger solid particles to break into smaller pieces to increase the rate of reaction
- Flour dust in flour mills and coal dusts in coal mines are hazardous because they can react very quickly, resulting in an explosion

V. Presence of catalysts

- Provides an alternative pathway which has a lower activation energy, providing a site for reaction
- Small amount of catalyst can give a large effect (hydrogen peroxide catalyst)
- Enzymes are organic catalysts which control the rate of biochemical reactions in living organisms: they are protein in nature and are destroyed by heat
- Misconception: “A catalyst speeds up the reaction but does not take part in it”
  - Catalysts take part in the reaction, but is itself not chemically reacted

Answering Technique Guide:

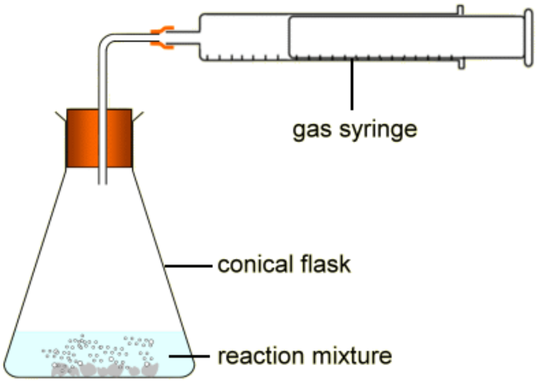
i. Increased X factor
ii. Increased frequency of effective collisions
iii. Between A and B reactant

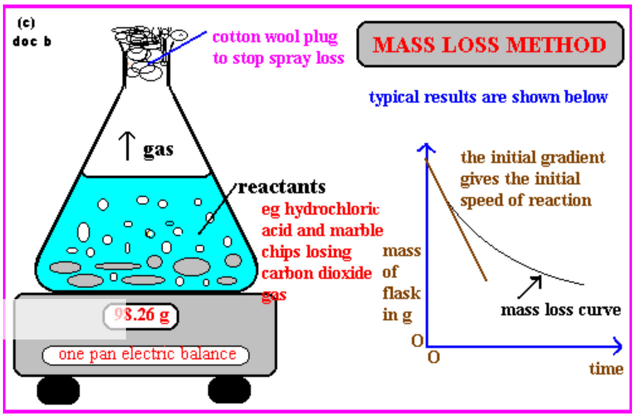
Role of Catalysts:

- Reactions occur more rapidly
  - Catalysts react with the reactants to form an intermediate compound (which then breaks back down into products and catalyst)
  - Lowers the activation energy since the particles can collide with catalyst as well as with each other (more particles have sufficient energy to get over the new energy barrier: this increases the rate of reaction)
- Catalysts are not sued or consumed during the reactions and do not appear as either the products or reactants of a reaction
- Importance in chemical industry: to produce products at an economical rate
  - Use of simple trial and error: most are transition metals (over 2500 catalysts tried for Haber process: production of ammonia)
- Heterogeneous catalyst: of another state compared to the reactants (most catalysts are heterogeneous and are in their solid form)
- Homogeneous: has the same state as the reactants
- Note: catalyst can become poisoned if certain molecules are preferentially adsorbed or permanently attached to the surface of the catalyst (reduces number of active sites and makes catalysts less effective): works the same way as enzyme inhibitors

Measuring the Rate of Reaction:

- Measuring the <u>decrease in the mass of the reaction mixture</u>
- Measuring the <u>increase in the volume of gas</u> formed





- The steeper the slope of the graph at any particular time, the faster the reaction at that time (usually graphs show that the reaction rate is fastest at the start of the reaction when the reactants are at their highest concentration)
- The steepness of the gradient decreases over time, hence the rate of reaction decreases
- When the gradient is zero, the reaction has stopped
- Assumption: one of the products must be a gas

Other Methods:

- Measuring the changes in the amount of a reactant or a product by titration (small portions of the reaction mixture is removed at fixed intervals and titrated to check the amount of reactant/product)
- Measuring the changes in intensity of the colour of a reaction mixture (coloured substance and intensity of colour can be measured using a colorimeter)
- Measuring changes in the pressure in a closed reaction vessel (suitable for reactions in a gaseous state, where the number of moles of gaseous products differ from the number of moles of the gaseous reactants (molar volume of gas)
- Measuring temperature change as the reaction proceeds

Experiments:

Thermite Reaction:

- Reaction between aluminium and a metal oxide (iron III)
- In a solid form (powder) it has no reaction
- A burning magnesium strip which releases quantities of energy provides energy to the reactants to overcome the activation energy (for a huge energy release)
- Temperatures can soar up to 3000 degrees Celsius

Lycopodium Powder:

- Heat energy is required for decomposition (heat provided by candle)
- High Activation Energy
- In a powder form to increase surface area for reaction (which produces pressure)